CHM113 General Chemistry I (Dr. Jen) Key

Extra Final Exam Practice Problems Chapters 11, 13, 14, 15, 16, 19

Chapter 11: IMFS!

1) As a gaseous element condenses, the atoms become closer together and they have more attraction for one another.

A) more separated, more B) more separated, less **C)** closer together, more
D) closer together, less E) larger, greater

2) Which one of the following exhibits dipole-dipole attraction between molecules?

A) XeF_4 **B)** AsH_3 C) CO_2 D) BCl_3 E) Cl_2

3) Of the following substances, only _____ has London dispersion forces as its only intermolecular force.

A) CH_3OH B) NH_3 C) H_2S **D)** CH_4 E) HCl

4) Which one of the following should have the ***lowest*** boiling point?

A) PH_3 B) H_2S C) HCl **D)** SiH_4 E) H_2O

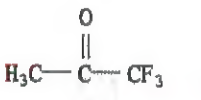
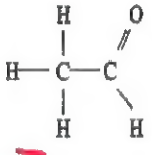
5) Of the following substances, _____ has the ***highest*** boiling point.

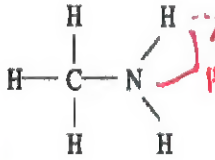
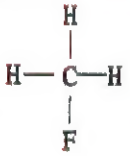
A) CH_3CH_2OH B) C_2H_6 C) N_2 D) F_2 **E)** $HOCH_2CH_2OH$

6) In which of the following molecules is hydrogen bonding likely to be the most significant component of the total intermolecular forces?

A) CH_4 B) $C_5H_{11}OH$ C) $C_6H_{13}NH_2$ **D)** CH_3OH E) CO_2

7) Which one of the following substances will have hydrogen bonding as one of its intermolecular forces?

A) B)  C) 

D)  E) 

8) What types of intermolecular forces exist between HI and H_2S?

A) dipole-dipole and ion-dipole
B) dispersion forces, dipole-dipole, and ion-dipole
C) dispersion forces, hydrogen bonding, dipole-dipole, and ion-dipole
D) dispersion forces and dipole-dipole
E) dispersion forces, dipole-dipole, and ion-dipole

Key

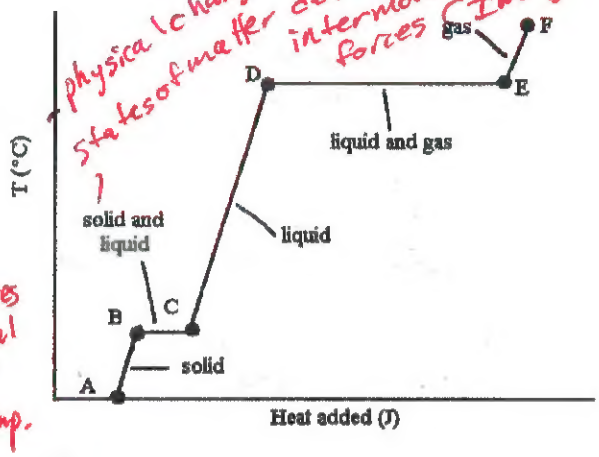
9) What type(s) of intermolecular forces exist between Cl_2 and CO_3^{-2}?

A) dispersion forces
B) dispersion forces and ion-dipole
C) dispersion forces, ion-dipole, and induced dipole - induced dipole
D) dispersion forces and ion-induced dipole
E) dispersion forces, ion-dipole, dipole-dipole, and ion-induced dipole

dispersion only
ion so can only have dispersion or ion-induced dispersion but no ion-dipole b/c Cl_2 has no dipole

10) The phase changes B → C and D → E are not associated with temperature increases because the heat energy is used up to:

A) increase distances between molecules
B) break intramolecular bonds
C) rearrange atoms within molecules
D) increase the velocity of molecules
E) increase the density of the sample



physical change
states of matter determined by intermolecular forces (IMFs)
requires chemical rxn
↑ temp.
is intensive property so won't Δ

11) Based on the following information, which compound has the strongest intermolecular forces?

most ΔH required

Substance	ΔH_{vap}(kJ/mol)
Argon (Ar)	6.3
Benzene (C_6H_6)	31.0
Ethanol (C_2H_5OH)	39.3
Water (H_2O)	40.8
Methane (CH_4)	9.2

A) Argon B) Benzene C) Ethanol D) Water E) Methane

Chapter 13: Solutions

H_2O

1) The dissolution of water in octane (C_8H_{18}) is prevented by _____.

solubility to mix & form a solution

A) London dispersion forces between octane molecules – weak IMF
B) hydrogen bonding between water molecules – strong IMF
C) dipole-dipole attraction between octane molecules – non-polar ∴ doesn't have this
D) ion-dipole attraction between water and octane molecules

no ions

 Key

2) Ammonium nitrate (NH_4NO_3) dissolves readily in water even though the dissolution is endothermic by 26.4 kJ/mol. The solution process is spontaneous because _____.

+ ΔH

$\Delta G = \Delta H - T\Delta S$
(−) (+)

A) the vapor pressure of the water decreases upon addition of the solute

B) osmotic properties predict this behavior ×

C) of the decrease in enthalpy upon addition of the solute → means exothermic

D) of the increase in enthalpy upon dissolution of this strong electrolyte → doesn't make spontaneous...

E) of the increase in disorder upon dissolution of this strong electrolyte

3) Which one of the following substances would be the ***most*** soluble in CCl_4? non-polar, dispersion = "like dissolves like"

A) CH_3CH_2OH B) H_2O C) NH_3 D) $C_{10}H_{22}$ E) NaCl

H-bond non-polar ionic

4) Which one of the following substances would be the ***least*** soluble in CCl_4? most unalike

A) CH_3CH_2OH B) H_2O C) NH_3 D) $C_{10}H_{22}$ E) NaCl

5) Which one of the following is ***most*** soluble in water? H-bonding, dipole-dipole or ion-dipole → most alike ∴ most H-bonding

A) CH_3OH B) $CH_3CH_2CH_2OH$ C) CH_3CH_2OH

largest % H-bonding

D) $CH_3CH_2CH_2CH_2OH$ E) $CH_3CH_2CH_2CH_2CH_2OH$

6) Which of the following statements is false?

A) Nonpolar liquids tend to be insoluble in polar liquids. ✓

B) The ~~weaker~~ stronger the attraction between the solute and solvent molecules, the greater the solubility.

C) Substances with similar intermolecular attractive forces tend to be soluble in one another. ✓

D) The solubility of a gas increases in direct proportion to its partial pressure above the solution. ✓ (though we didn't cover this) it makes sense that putting pressure should force a gas into a liquid to make a soln

E) The solubility of gases in water decreases with increasing temperature. ✓

Key

7) Which of the following choices has the compounds correctly arranged in order of increasing solubility in water? (least soluble to most soluble)

A) CCl_4 < $CHCl_3$ < $NaNO_3$ (non-polar < dipole-dipole < ionic)

B) CH_3OH < CH_4 < LiF (would be best → H-bonding < disp. < ionic)

C) CH_4 < $NaNO_3$ < $CHCl_3$ (non-polar < ionic < dipole-dipole)

D) LiF < $NaNO_3$ < $CHCl_3$ (ionic < ionic < dipole-dipole)

E) CH_3OH < Cl_4 < $CHCl_3$ (H-bonding < nonpolar < dipole-dipole)

Chapter 14 – Kinetics!

1) At elevated temperatures, dinitrogen pentoxide decomposes to nitrogen dioxide and oxygen:

$$2N_2O_5(g) \rightarrow 4\,NO_2(g) + O_2(g)$$

$x = ?$; 2.2×10^{-4}

$$2.2\times10^{-4} \times \frac{2\,mol\,N_2O_5}{1\,mol\,O_2} = 4.4\times10^{-4}\ M/s$$

When the rate of formation of O_2 is 2.2×10^{-4} M/s, the rate of decomposition of N_2O_5 is _____ M/s.

A) 1.1×10^{-4} B) 2.2×10^{-4} C) 2.8×10^{-4} D) 4.4×10^{-4}

E) 5.5×10^{-4}

M/s ⊥ M/s M $M = \frac{mol}{L}$ so $\frac{1}{M\cdot s} = \frac{L}{mol\cdot s}$

2) The rate law of a reaction is rate = k[D][X]. The units of the rate constant are _____.

A) mol $L^{-1}s^{-1}$ B) L $mol^{-1}s^{-1}$ C) $mol^2\ L^{-2}s^{-1}$

D) mol $L^{-1}s^{-2}$

3) *The data in the table below were obtained for the reaction:*

A + B → P

Experiment Number	[A] (M)	[B] (M)	Initial Rate (M/s)
1	0.273	0.763	2.83
2	0.273	1.526	2.83
3	0.819	0.763	25.47

3× ; 2× ; 9× no Δ ∴ $[B]^0$

$\therefore [3A]^x = 9$ so $[A]^2$

$x = 2$ so rate $= k[A]^2$

The rate law for this reaction is rate = _____.

A) k[A][B] B) k[P] C) $k[A]^2[B]$ D) $k[A]^2[B]^2$ E) $k[A]^2$

4) The magnitude of the rate constant for the above reaction is: ___.

A) 38.0 B) 0.278 C) 13.2 D) 42.0 E) 2.21

rate $= k[A]^2$

$$k = \frac{rate}{[A]^2} = \frac{2.83}{[0.273]^2} = 37.97 = 38.0$$

 Key

5) *The data in the table below were obtained for the reaction:*

$$2\ ClO_2\ (aq) + 2\ OH^-\ (aq) \rightarrow ClO_3^-\ (aq) + ClO_2^-\ (aq) + H_2O\ (l)$$

Experiment Number	$[ClO_2]$ (M)	$[OH^-]$ (M)	Initial Rate (M/s)
1	0.060	0.030	0.0248
2	0.020	0.030	0.00276
3	0.020	0.090	0.00828

$(3)^x = 9$ so $[ClO_2]^2$

$(3)^x = 3$ so $[OH^-]^1$

a) What is the order of the reaction with respect to ClO_2?

A) 1 B) 0 C) 2 D) 3 E) 4

b) What is the order of the reaction with respect to OH^-?

A) 0 B) 1 C) 2 D) 3 E) 4

c) What is the overall order of the reaction? 2 + 1 = 3

A) 4 B) 0 C) 1 D) 2 E) 3

sum

rate = $k[OH^-]^1[ClO_2]^2$

d) What is the magnitude of the rate constant for the reaction?

A) 1.15×10^4 B) 4.6 C) 230 D) 115 E) 713

$$k = \frac{\text{rate}}{[OH^-][ClO_2]^2} = \frac{(0.0248 M/s)}{(0.030 M)(0.060)^2} = 229.6 = 230$$

2 sig figs

6) The decomposition of N_2O_5 in solution in carbon tetrachloride proceeds via the reaction:

$2\ N_2O_5\ (soln) \rightarrow 4\ NO_2\ (soln) + O_2\ (soln)$

The reaction is first order and has a rate constant of $4.82 \times 10^{-3}\ s^{-1}$ at 64°C. The rate law for the reaction is:

only 1 reactant to rxn order of 1

rate = $k[N_2O_5]^1$

rate =

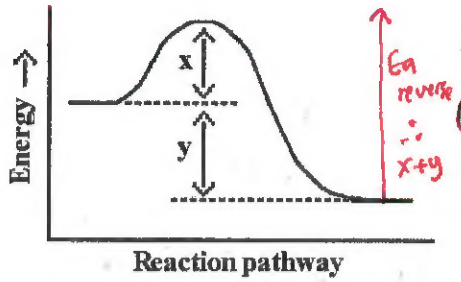
A) $k[N_2O_5]^2$ ← 2nd order B) $k\frac{[NO_2]^4[O_2]}{[N_2O_5]^2}$ C) $k[N_2O_5]$ D) $k\frac{[N_2O_5]^2}{[NO_2]^4[O_2]}$

E) $2k[N_2O_5]$?

no pdts in rate laws

Key

7) Which energy difference in the energy profile below corresponds to the activation energy for the reverse reaction?



Ea reverse ∴ x+y

A) x
B) y
C) x + y
D) x - y
E) y - x

8) In general, as temperature goes up, reaction rate _____.

A) goes up if the reaction is exothermic

B) goes up if the reaction is endothermic

C) goes up regardless of whether the reaction is exothermic or endothermic

D) stays the same regardless of whether the reaction is exothermic or endothermic

E) stays the same if the reaction is first order

b/c molecules can collide more frequently

9) The mechanism for formation of the product X is:

A + B → C + D (slow) — D formed
B + D → X (fast) — D consumed

∴ not in overall rxn ∴ intermediate

The intermediate reactant in the reaction is _____.

A) A B) B C) C D) D E) X

10) For the elementary reaction: $NO_3 + CO \rightarrow NO_2 + CO_2$ the molecularity of the reaction is 1+1 = 2, and the rate law is rate = $k[NO_3][CO]$.

∴ coefficients = rxn order

A) 2, k[NO_3][CO] B) 4, k[NO_3][CO][NO_2][CO_2] C) 2, k[NO_2][CO_2]

D) 2, k[NO_3][CO]/[NO_2][CO_2] E) 4, k[NO_2][CO_2]/[NO_3][CO]

Key

11) A possible mechanism for the overall reaction: **Br_2 (g) + 2NO (g) → 2NOBr (g)**
Is:

$NO\ (g) + Br_2\ (g) \underset{k_{-1}}{\overset{k_1}{\rightleftharpoons}} NOBr_2\ (g)$ (fast)

$NOBr_2\ (g) + NO\ (g) \xrightarrow{k_2} 2NOBr$ (slow)

$rate_{for} = k_1[NO][Br_2] = k_{-1}[NOBr_2] = rate_{rev}$

$[NOBr_2] = \frac{k_1[NO][Br_2]}{k_{-1}}$

$rate = k_2[NOBr_2][NO]$

R

intermediate ∴ can't be in the rate law

The rate law for formation of NOBr based on this mechanism is rate = _____. so

A) $k_1[NO]1/2$ B) $k_1[Br_2]1/2$ C) $(k_2k_1/k^{-1})[NO]^2[Br_2]$ D) $(k_1/k^{-1})^2[NO]^2$ E) $(k_2k_1/k^{-1})[NO]^2[Br_2]^2$

$rate = \left(\frac{k_1 \cdot k_2}{k_{-1}}\right)[NO][Br_2]$ (same reactant)

so

$rate = \left(\frac{k_1k_2}{k_{-1}}\right)[Br_2][NO]^2$

12) Which of the following is true?

A) If we know that a reaction is an elementary reaction, then we know its rate law. ✔

B) The rate-determining step of a reaction is the rate of the ~~fastest~~ slowest elementary step of its mechanism.

C) Since intermediate compounds can be formed, the chemical equations for the elementary reactions in a multistep mechanism do not always have to add to give the chemical equation of the overall process. – the steps must add to the overall process

D) In a reaction mechanism, an intermediate is identical to an activated complex. X

transition state

13) Consider the following reaction: **3A → 2B** The average rate of appearance of B is given by Δ[B]/Δt. Comparing the rate of appearance of B and the rate of disappearance of A, we get Δ[B]/Δt = _____ × (-Δ[A]/Δt).

A) -2/3 B) +2/3 C) -3/2 D) +1 E) +3/2

$ave\ rate = -\left(\frac{1}{3}\right)\frac{\Delta[A]}{\Delta t} = +\left(\frac{1}{2}\right)\frac{\Delta[B]}{\Delta t}$ so $\frac{\Delta[B]}{\Delta t} = \left(\frac{2}{3}\right)\frac{-\Delta[A]}{\Delta t}$

$aA \rightarrow bB$

$-\frac{1}{a}\frac{\Delta[A]}{\Delta t} = +\frac{1}{b}\left(\frac{\Delta[B]}{\Delta t}\right)$

Key

Chapter 15 – Equilibrium

1) Which one of the following will change the value of an equilibrium constant?

(A) changing temperature ✓

B) adding other substances that do not react with any of the species involved in the equilibrium ✗

C) varying the initial concentrations of reactants ← no @ equil yet ∴ not in equilibrium expression

D) varying the initial concentrations of products

E) changing the volume of the reaction vessel ← matters when gases but will shift equilibrium not change the constant

2) Which of the following expressions is the correct equilibrium-constant expression for the equilibrium between dinitrogen tetroxide and nitrogen dioxide? $\mathbf{N_2O_4\ (g) \rightleftharpoons 2NO_2\ (g)}$

$K_{eq} = \frac{[\text{prods}]^a_{eq}}{[\text{reactants}]^b_{eq}} = \frac{[NO_2]^2}{[N_2O_4]}$

A) $\frac{[NO_2]}{[N_2O_4]}$ ✗ (B) $\frac{[NO_2]^2}{[N_2O_4]}$ ✓ C) $\frac{[NO_2]}{[N_2O_4]}$ ✗ D) $[NO_2]^2[N_2O_4]$ ✗

3) The equilibrium constant for reaction 1 is K. The equilibrium constant for reaction 2 is _____.

(1) $\mathbf{SO_2\ (g) + (1/2)\ O_2\ (g) \rightleftharpoons SO_3\ (g)}$ $K_{eq} = K$

(2) $\mathbf{2\ SO_3\ (g) \rightleftharpoons 2\ SO_2\ (g) + O_2\ (g)}$ $K_{eq} = \frac{1}{K^2}$

reversed & multiple = 2

A) K^2 B) 2 K C) ½ K (D) $1/K^2$ E) $-K^2$

4) The value of K_{eq} for the following reaction is 0.25:

$\mathbf{SO_2\ (g) + NO_2\ (g) \rightleftharpoons SO_3\ (g) + NO\ (g)}$ $K_{eq} = 0.25$

The value of K_{eq} at the same temperature for the reaction below is _____.

$\mathbf{2\ SO_2\ (g) + 2\ NO_2\ (g) \rightleftharpoons 2\ SO_3\ (g) + 2\ NO\ (g)}$

not reversed multiple = 2

$K = (0.25)^2 = 0.0625 = 0.06$ 2 sig. figs.

A) 0.50 (B) 0.063 C) 0.12 D) 0.25 E) 16

5) The equilibrium constant for the gas phase reaction: $\mathbf{N_2\ (g) + 3\ H_2\ (g) \rightleftharpoons 2\ NH_3\ (g)}$ is $K_{eq} = 4.34 \times 10^{-3}$ at 300 °C. At equilibrium, _____.

reactants favored

A) products predominate ✗ (B) reactants predominate

C) roughly equal amounts of products and reactants are present would be ≈ 1

D) only products are present ✗ E) only reactants are present ✗

Key

6) Of the following equilibria, only _____ will shift to the left in response to a decrease in volume. matters for gases, if ↓V then favor side w/ fewer moles of gas (b/c less room for the molecules to move around in)

A) H_2 (g) + Cl_2 (g) ⇌ 2 HCl (g) – same # gas moles ∴ won't shift

B) 2 SO_3 (g) ⇌ 2 SO_2 (g) + O_2 (g) fewer moles

C) N_2 (g) + 3 H_2 (g) ⇌ 2 NH_3 (g) 2 fewer

D) 4 Fe (s) + 3 O_2 (g) ⇌ 2 Fe_2O_3 (s) fewer moles gas (b/c no gas)

E) 2 HI (g) ⇌ H_2 (g) + I_2 (g) same ∴ no shift

7) The reaction below is exothermic: **2 SO_2 (g) + O_2 (g) ⇌ 2 SO_3 (g) + heat** (↑ shift right = desired)

_____ will result in an increase in the number of moles of SO_3 (g) in the reaction container.

A) increasing the pressure (favors less gas →) ✓ B) decreasing the pressure (favors more gas ←)

C) increasing the temperature (← away from increase) D) removing some oxygen (← toward a decrease)

E) increasing the volume of the container – shift towards more moles gas ←

8) Consider the following reaction at equilibrium: (↑ ← ↑ ↓) **2 NH_3 (g) ⇌ N_2 (g) + 3 H_2 (g)**

$\Delta H° = +92.4$ kJ that adding N_2 (g) to the system at equilibrium will result in _____.

A) a decrease in the concentration of NH_3 (g) ✗

B) a decrease in the concentration of H_2 (g) ✓

C) an increase in the value of the equilibrium constant ✗ won't change

D) a lower concentration of N_2 ✗

E) removal of all of the H_2 (g) ✗ → then wouldn't be under ⇌ conditions

9) Consider the following reaction at equilibrium: (→ ↑) **C (s) + H_2O (g) ⇌ CO (g) + H_2 (g)**

Which of the following conditions will increase the concentration of CO?

A) decreasing the concentration of H_2O (g) ↓ (toward ← decrease) ✗

B) removing H_2O (g) from the system ↓ ← ✗

C) decreasing the volume of the reaction vessel – favor side w/ fewer moles gas ← ✗

D) increasing the amount of carbon in the system ↑ → ✓

Key

10) The value of K_{eq} for the equilibrium: **H_2 (g) + I_2 (g) ⇌ 2 HI (g)** is 794 at 25°C. At this temperature, what is the value of K_{eq} for the equilibrium:

HI (g) ⇌ 1/2 H_2 (g) + 1/2 I_2 (g) ?

reverse & multiple = ½ so $K = \frac{1}{(794)^{1/2}} = 0.03548$

A) 1588 B) 28 C) 397 D) 0.0355 E) 0.0013

14) Consider the following chemical reaction: **CO (g) + 2 H_2 (g) ⇌ CH_3OH(g)**
At equilibrium in a particular experiment, the concentrations of CO and H_2 were 0.15 M and 0.36 M, respectively. What is the equilibrium concentration of CH_3OH? The value of K_{eq} for this reaction is 14.5 at the temperature of the experiment.

$K_c = \frac{[CH_3OH]_{eq}}{[CO]_{eq}[H_2]_{eq}^2}$

A) 14.5 B) 7.61 × 10^{-3} C) 2.82 × 10^{-1}
D) 3.72 × 10^{-3} E) 1.34 × 10^{-3}

$K_c = 14.5$ $[CO]_{eq} = 0.15M$ $[H_2]_{eq} = 0.36M$

$[CH_3OH]_{eq} = K_c[CO]_{eq}[H_2]_{eq}^2 = (14.5)(0.15)(0.36)^2$

$K_c = 0.28188 = 2.82 \times 10^{-1}$

15) In the coal-gasification process, carbon monoxide is converted to carbon dioxide via the following reaction: **CO (g) + H_2O (g) ⇌ CO_2 (g) + H_2 (g)** In an experiment, 0.35 mol of CO and 0.40 mol of H_2O were placed in a 1.00-L reaction vessel. At equilibrium, there were 0.16 mol of CO remaining. K_{eq} at the temperature of the experiment is _____.

$[CO]_i = \frac{0.35 mol}{1.00 L} = 0.35M$ $[H_2O]_i = \frac{0.40 mol}{1.00 L} = 0.40M$

A) 5.5 B) 0.75 C) 0.93 D) 1.1 E) 1.0

$[CO]_{eq} = \frac{0.16 mol}{1.00 L} = 0.16M$

	CO + H_2O ⇌	CO_2 +	H_2
I (M)	0.35 0.40	0	0
C (M)	−0.19 −0.19	+0.19	+0.19
E (M)	0.16 0.21	0.19	0.19

* All 1:1 molar ratio ∴ change is the same

$K_{eq} = \frac{[CO_2][H_2]_{eq}}{[CO]_{eq}[H_2O]_{eq}} = \frac{(.19)(.19)}{(.16)(.21)} = 1.074$

$K_{eq} = 1.1$

16) Nitrosyl bromide decomposes according to the following equation.
2 NOBr (g) ⇌ 2 NO (g) + Br_2 (g) A sample of NOBr (0.64 mol) was placed in a 1.00-L flask containing no NO or Br_2. At equilibrium the flask contained 0.36 mol of NOBr. How many moles of NO and Br_2, respectively, are in the flask at equilibrium?

A) .28,.28 B) .36,.18 C) .28,.14 D) .14,.23 E) .36,.36

$[NOBr]_i = \frac{0.64 mol}{1.00 L} = 0.64M$

$[NOBr]_{eq} = \frac{0.36 mol}{1.00 L} = 0.36M$

$\Delta[NOBr] = \frac{0.28 mol NOBr}{L sol}$

NO: $\times \frac{2 mol NO}{2 mol NOBr} \times 1L = 0.28 mol$

Br_2: $\times \frac{1 mol Br_2}{2 mol NOBr} \times 1L = 0.14 mol$

Key

Chapter 16 Acid-Base Equilibria

1) Which one of the following is a BrØnsted -Lowry acid? donates H^+

A) $(CH_3)_3NH^+$ B) CH_3COOH C) HF D) HNO_2

E) all of the above

2) Which one of the following is a BrØnsted -Lowry base? accepts H^+

A) $(CH_3)_3N$: B) CH_3COOH C) HF D) HNO_2

E) none of the above

3) Which solution below has the highest concentration of hydroxide ions?

A) pH = 3.21 B) pH = 12.6 C) pH = 7.93 D) pH = 9.82

E) pH = 7.00

4) Of the following acids, _____ is a strong acid.

A) HNO_2 B) H_2CO_3 C) $HBrO_3$ D) HClO E) HF

5) Of the following acids, _____ is not a strong acid.

A) HNO_2 B) H_2SO_4 C) HNO_3 D) $HClO_4$ E) HCl

Acid	K_a
HOAc	1.8×10^{-5}
$HCHO_2$	1.8×10^{-4}
HClO	3.0×10^{-8}
HF	6.8×10^{-4}

6) Of the acids in the table, _____ is the strongest acid.

A) HOAc B) $HCHO_2$ C) HClO

D) HF E) HOAc and $HCHO_2$

$HA \rightleftharpoons H^+ + A^-$ (ionization) $K_a = \frac{[H^+][A^-]}{[HA]}$

7) In which of the following aqueous solutions does the weak acid exhibit the highest percentage ionization? largest K_a

A) 0.01 M $HC_2H_3O_2$ ($K_a = 1.8 \times 10^{-5}$) B) 0.01 M HNO_2 ($K_a = 4.5 \times 10^{-4}$)

C) 0.01 M HF ($K_a = 6.8 \times 10^{-4}$) D) 0.01 M HClO ($K_a = 3.0 \times 10^{-8}$)

E) These will all exhibit the same percentage ionization.

8) HA is a weak acid. Which equilibrium corresponds to the equilibrium constant K_b for A^-?

$A^-_{(aq)} + H_2O_{(l)} \rightleftharpoons HA_{(aq)} + OH^-_{(aq)}$

A) $HA\ (aq) + H_2O\ (l) \rightleftharpoons H_2A^+\ (aq) + OH^-\ (aq)$

B) $A^-\ (aq) + H_3O^+\ (aq) \rightleftharpoons HA\ (aq) + H_2O\ (l)$

C) $HA\ (aq) + OH^-\ (aq) \rightleftharpoons H_2O\ (l) + H^+\ (aq)$

D) $A^-\ (aq) + H_2O\ (l) \rightleftharpoons HA\ (aq) + OH^-\ (aq)$ E) $A^-\ (aq) + OH^-\ (aq) \rightleftharpoons HOA^{2-}\ (aq)$

Key

9) Using the data in the table,

$(K_a)(K_b)=K_w$

Acid	K_a
HOAc	1.8×10^{-5}
$HC_7H_5O_2$	6.3×10^{-5}
HNO_2	4.5×10^{-4}
HF	6.8×10^{-4}

smallest

a) which of the conjugate bases below is the strongest base?

(A) OAc^- B) $C_7H_5O_2^-$ C) NO_2^- D) F^- E) OAc^- and $C_7H_5O_2^-$

↳ will be weakest acid

b) which of the conjugate bases below is the weakest base?

strongest acid ∴ largest K_a

A) OAc^- B) $C_7H_5O_2^-$ C) NO_2^- (D) F^- E) OAc^- and $C_7H_5O_2^-$

10) Which of the following ions will act as a weak base in water?

come from weak acids

A) HS^- ✓ H_2S weak acid B) F^- ✓ HF weak acid C) NO_2^- ✓ HNO_2 weak acid D) ClO^- ✓ HClO weak acid

(E) All of the above will act as a weak base in water.

11) a) Which of the following aqueous solutions has the highest $[OH^-]$?

basic or @ least not least acidic

A) a solution with a pH of 3.0 ∴ acidic B) a 1×10^{-4} M solution of HNO_3 strong acid ∴ pH < 7

C) a solution with a pOH of 12.0 (14.0 − 12.0 = pH = 2 ∴ acidic) (D) pure water pH = 7

E) a 1×10^{-3} M solution of NH_4Cl negligible base

NH_4^+ weak acid so pH < 7

b) Which of the following aqueous solutions has the lowest $[OH^-]$?

↳ most acidic

A) a solution with a pH of 3.0 B) a 1×10^{-4} M solution of HNO_3 strong acid ∴

(C) a solution with a pOH of 12.0 (pH = 2.0) D) pure water pH = 7.00

$pH = -\log[1 \times 10^{-4}]$

$pH = 4.0$

E) a 1×10^{-3} M solution of NH_4Cl

weak acid ∴ pH not as low as

12) Of the following substances, an aqueous solution of _____ will form basic solutions.

H_2S = weak acid

HF = weak acid ∴ weak base = $F^\ominus$

NaHS **$Cu(NO_3)_2$** **$KHCO_3$** **NaF**

NaHS: weak base

$Cu(NO_3)_2$: negligible base ble from strong acid HNO_3

$KHCO_3$: H_2CO_3 carbonic acid = weak acid ∴ HCO_3^- = weak base

A) NaHS, $Cu(NO_3)_2$ B) $KHCO_3$, NaHS C) NaF only

D) NaF, $KHCO_3$ (E) NaHS, $KHCO_3$ and NaF

Key

$(K_a)(K_b) = K_w = 1.00\times10^{-14}$

13) Of the compounds below, a 0.1 M aqueous solution of _____ will have the highest pH.

most basic ∴ smallest K_a or largest K_b

(A) KCN, K_a of HCN = 4.0×10^{-10} (base; $K_b = 2.5\times10^{-5}$)

B) NH_4NO_3, K_b of NH_3 = 1.8×10^{-5} (acidic soln; negligible base)

C) NaOAc, K_a of HOAc = 1.8×10^{-5} (base; $K_b = 5.5\times10^{-10}$)

D) NaClO, K_a of HClO = 3.2×10^{-8} (weak base; $K_b = 3.125\times10^{-7}$)

E) NaHS, K_b of HS^- = 1.8×10^{-7} (base)

14) Of the following, which is the strongest acid?

A) HIO (B) HIO_4 C) HIO_2 D) HIO_3 E) They are all the same.

largest # of oxygens

15) Which of the following acids will be the strongest?

(A) H_2SO_4 B) HSO_4^- ✗ C) H_2SO_3 D) H_2SeO_4 E) HSO_3^- ✗

fewer Oxygens; less EN ∴ less polarization

16) What is the conjugate base of OH^-? (remove H^+) → $O^{2-} + H^+$

A) O_2 B) O^- C) H_2O (D) O^{2-} E) H_3O^+

17) What is the pOH of an aqueous solution at 25.0 °C that contains 3.98×10^{-9} M hydronium ion?

$[H^+][OH^-] = 1\times10^{-14}$; $[OH^-] = \frac{1.00\times10^{-14}}{3.98\times10^{-9}} = 2.51\times10^{-6}$; $pOH = -\log(2.51\times10^{-6}) = 5.600$

$[H_3O^+] \approx [H^+] = 3.98\times10^{-9}$ M; $pH = -\log(3.98\times10^{-9}) = 8.400$; $pH + pOH = 14.00$; $14.00 - 8.400 = 5.600$

A) 8.400 (B) 5.600 C) 9.000 D) 3.980 E) 7.000

either approach works

18) Calculate the concentration (in M) of hydronium ions in a solution at 25.0°C with a pOH of 4.223.

$14.00 - pOH = pH$; $14.00 - 4.223 = 9.777$; $pH = -\log[H^+]$; $[H^+] = 10^{-pH} = 10^{-9.777} = 1.67\times10^{-10}$

$[OH^-] = 10^{-pOH}$; $[OH^-] = 10^{-4.223} = 5.98\times10^{-5}$; $[H^+] = \frac{1.0\times10^{-14}}{[OH^-]} = \frac{1.0\times10^{-14}}{5.98\times10^{-5}}$; $[H^+] = 1.67\times10^{-10}$

A) 5.98×10^{-5} (B) 1.67×10^{-10} C) 1.67×10^{4}

D) 5.99×10^{-19} E) 1.00×10^{-7}

either way works!

19) An aqueous solution contains 0.100 M NaOH at 25.0°C. The pH of the solution is _____.

A) 0.100 B) 1.00 (C) 13.00 D) 7.00 E) -1.00

$pOH = -\log[OH^-] = -\log(0.100) = 1.00$; $pH = 14.000 - pOH = 14.00 - 1.00$; $pH = 13.00$

$[H^+] = \frac{1.0\times10^{-14}}{0.100} = 1.00\times10^{-13}$; $pH = -\log[H^+] = -\log(1.00\times10^{-13}) = 13.00$

20) The pH of a 0.55 M aqueous solution of hypobromous acid, HBrO, at 25.0°C is 4.48. What is the value of K_a for HBrO?

$K_a = \frac{[H^+]_{eq}[BrO^-]_{eq}}{[HBrO]_{eq}}$

$HBrO_{(aq)} \rightleftharpoons H^+_{(aq)} + BrO^-_{(aq)}$

	HBrO	H^+	BrO^-
I	0.55	0	0
C	-3.3×10^{-5}	$+3.3\times10^{-5}$	$+3.3\times10^{-5}$
E	0.55	3.3×10^{-5}	3.3×10^{-5}

(A) 2.0×10^{-9} B) 1.1×10^{-9} C) 6.0×10^{-5}

D) 3.3×10^{-5} E) 3.0×10^{4}

$pH = -\log[H^+]_{eq}$; $[H^+]_{eq} = 10^{-pH} = 10^{-4.48} = 3.3\times10^{-5}$

$K_a = \frac{[3.3\times10^{-5}][3.3\times10^{-5}]}{[0.55]} = 1.98\times10^{-9} = 2.0\times10^{-9}$

Key

$(K_a)(K_b) = K_w = 1.00 \times 10^{-14}$

21) The acid-dissociation constant, K_a, for gallic acid is 4.57×10^{-3}. What is the base-dissociation constant, K_b, for the gallate ion?

$K_b = \frac{1.00 \times 10^{-14}}{(4.57 \times 10^{-3})} = 2.19 \times 10^{-12}$

A) 4.57×10^{-3} B) 2.19×10^{-12} C) 5.43×10^{-5}
D) 7.81×10^{-6} E) 2.19×10^{2}

22) The K_a for HCN is 4.9×10^{-10}. What is the value of K_b for CN^-?

A) 2.0×10^{-5} B) 4.0×10^{-6} C) 4.9×10^{4} D) 4.9×10^{-24}
E) 2.0×10^{9}

$K_b = \frac{1.00 \times 10^{-14}}{4.9 \times 10^{-10}} = 2.0 \times 10^{-5}$

23) What is the pOH of a 0.035 M solution of barium hydroxide?

$Ba(OH)_2 \rightarrow Ba^{2+}_{(aq)} + 2OH^-_{(aq)}$ strong base

$[OH^-] = 0.035\ mol\ Ba(OH)_2 / 1\ L\ soln \times \frac{2\ mol\ OH^-}{1\ mol\ Ba(OH)_2} = 0.0$ M

A) 12.85 B) 12.54 C) 1.46 D) 1.15 E) 10.41

$pOH = -\log[OH^-] = -\log(0.070) = 1.15 = pOH$

Chapter 19 – Thermodynamics

1) Which one of the following is always positive when a spontaneous process occurs?

A) ΔS_{system} B) $\Delta S_{surroundings}$ C) $\Delta S_{universe}$
D) $\Delta H_{universe}$ E) $\Delta H_{surroundings}$

2) Which one of the following processes produces a decrease of the entropy of the system?

↳ more order

A) dissolving sodium chloride in water – more disorder ↑S
B) melting of naphthalene s → l more disorder ↑S
C) dissolving oxygen in water – more ordered ↓S
D) boiling of alcohol l → g more disorder ↑S
E) explosion of nitroglycerine – more disorder ↑S

3) ΔS is positive for the reaction _____.

A) $2\ NO\ (g) + O_2\ (g) \rightarrow 2\ NO_2\ (g)$ ↓S (3 → 2)
B) $2\ N_2\ (g) + 3\ H_2\ (g) \rightarrow 2\ NH_3\ (g)$ ↓S (5 → 2)
C) $C_3H_8\ (g) + 5\ O_2\ (g) \rightarrow 3\ CO_2\ (g) + 4\ H_2O\ (g)$ ↑S (6 → 7)
D) $Mg\ (s) + Cl_2\ (g) \rightarrow MgCl_2\ (s)$ ↓S
E) $C_2H_4\ (g) + H_2\ (g) \rightarrow C_2H_6\ (g)$ ↓S (2 → 1)

Key

4) ΔS is positive for the reaction _____.

A) $2\ Ca\ (s) + O_2\ (g) \rightarrow 2\ CaO\ (s)$ ↓S

B) $2\ KClO_3\ (s) \rightarrow 2\ KCl\ (s) + 3\ O_2\ (g)$ ↑S

C) $HCl\ (g) + NH_3\ (g) \rightarrow NH_4Cl\ (s)$ ↓S

D) $Pb^{+2}\ (aq) + 2\ Cl^-\ (aq) \rightarrow PbCl_2\ (s)$ ↓S

E) $CO_2\ (g) \rightarrow CO_2\ (s)$ ↓S

5) Of the following, the entropy of _____ is the largest.

- g > l > s
- highest # atoms
- highest M

A) HCl (l) B) HCl (s) C) HCl (g) D) HBr (g) E) HI (g)

6) For an isothermal process, the entropy change of the surroundings is given by the equation: (same temp.)

A) $\Delta S = q_{sys}\ T$ B) $\Delta S = -q_{sys}\ T$ C) $\Delta S = q\ \ln T$ D) $\Delta S = -q\ \ln T$
E) $\Delta S = -q_{sys}/T$

7) A reaction that is not spontaneous at low temperature (ΔH) can become spontaneous at high temperature if ΔH is _____ and ΔS is _____.

ΔG° = neg ← large # will make
$= \Delta H^\circ - T\Delta S^\circ$
(+) (+)

A) +, + B) -, - C) +, - D) -, + E) +, 0

↖ can't get here yet

8) For the reaction: **$2\ C_4H_{10}\ (g) + 13\ O_2\ (g) \rightarrow 8\ CO_2\ (g) + 10\ H_2O\ (g)$**
ΔH° is -125 kJ/mol and ΔS° is +253 J/K · mol. This reaction is _____.

$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$
(-) (-) (+)

A) spontaneous at all temperatures
B) spontaneous only at high temperature
C) spontaneous only at low temperature
D) nonspontaneous at all temperatures
E) unable to determine without more information

Key

9) Consider the reaction: **Ag^+ (aq) + Cl^- (aq) → AgCl (s)**
Given the following table of thermodynamic data, determine the temperature (in K) above which the reaction is nonspontaneous under standard conditions.

Substance	ΔH_f° (kJ/mol)	S° (J/mol · K)
Ag^+ (aq)	105.90	73.93
Cl^- (aq)	-167.2	56.5
AgCl (s)	-127.0	96.11

Te which $\Delta G^\circ > 0$ positive

A) 1230 B) 150 C) 432 D) 133 E) 1914

$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$

$0 = \Delta H^\circ - T\Delta S^\circ$

$T = \frac{\Delta H^\circ}{\Delta S^\circ}$

$\Delta H^\circ = \sum n H_f^\circ \text{ pdts} - \sum m H_f^\circ \text{ reactants}$

$\Delta H^\circ = [1 H_f^\circ_{AgCl}] - [(1 H_f^\circ_{Ag^+}) + (1 H_f^\circ_{Cl^-})]$

$= (-127.0\text{kJ}) - [(105.90\text{kJ}) + (-167.2\text{kJ})]$

$= -127.0\text{kJ} - (-61.3\text{kJ})$

$\Delta H^\circ = -65.7\text{ kJ}$

$\Delta S^\circ = \sum n S^\circ \text{ pdts} - \sum m S^\circ \text{ reactants}$

$= [1 S^\circ_{AgCl}] - [(1 S^\circ_{Ag^+}) + (1 S^\circ_{Cl^-})]$

$= [96.11\text{J/K}] - [(73.93\text{J/K}) + (56.5\text{J/K})]$

$= -34.3\text{J/K} \times \frac{1\text{kJ}}{1000\text{J}} = -0.0343\text{ kJ/K}$

$T = \frac{(-65.7\text{ kJ})}{(-0.0343\text{kJ/K})} = 1914\text{ K}$

10) If ΔG° for a reaction is greater than zero, then _____.

A) K = 0 B) K = 1 C) K > 1 (D) K < 1

E) More information is needed.

$\Delta G^\circ = -RT\ln K$ (+)(+) has to be <1 to get pos. #

ΔG° positive → non-spontaneous

makes sense b/c reactants would be favored

Key

Use the table below to answer the questions that follow.

Thermodynamic Quantities for Selected Substances at 298.15 K (25°C)

Substance	$\Delta H°_f$ (kJ/mol)	$\Delta G°_f$ (kJ/mol)	S (J/K-mol)
Carbon			
C (s, diamond)	1.88	2.84	2.43
C (s, graphite)	0	0	5.69
C_2H_2 (g)	226.7	209.2	200.8
C_2H_4 (g)	52.30	68.11	219.4
C_2H_6 (g)	-84.68	-32.89	229.5
CO (g)	-110.5	-137.2	197.9
CO_2 (g)	-393.5	-394.4	213.6
Hydrogen			
H_2(g)	0	0	130.58
Oxygen			
O_2 (g)	0	0	205.0
H_2O (l)	-285.83	-237.13	69.91

11) The combustion of acetylene in the presence of excess oxygen yields carbon dioxide and water: **$2\ C_2H_2\ (g) + 5\ O_2\ (g) \rightarrow 4\ CO_2\ (g) + 2\ H_2O\ (l)$**

The value of $\Delta S°$ for this reaction is _____ J/K· mol.

A) +689.3 B) +122.3 C) +432.4 D) -122.3 E) -432.4

$\Delta S° = \sum n S°_{pdts} - \sum m S°_{reactants}$

$= [(4 S°_{CO_2}) + (2 S°_{H_2O})] - [(2 S°_{C_2H_2}) + (5 S°_{O_2})]$

$= [(4 \times \frac{213.6\,J}{mol \cdot K}) + (2 \times \frac{69.91\,J}{mol \cdot K})] - [(2 \times \frac{200.8\,J}{mol \cdot K}) + (5 \times \frac{205.0\,J}{mol \cdot K})]$

$= [\frac{854.4\,J}{mol \cdot K} + \frac{139.8\,J}{mol \cdot K}] - [\frac{401.6\,J}{mol \cdot K} + \frac{1025\,J}{mol \cdot K}]$

$= [\frac{994.2\,J}{mol \cdot K}] - [\frac{1426.6\,J}{mol \cdot K}]$

$\Delta S° = -432.4\,J/mol \cdot K$

Key

12) The value of ΔG° at 25°C for the formation of $POCl_3$ from its constituent elements,
P_2 (g) + O_2 (g) + 3 Cl_2 (g) → 2 $POCl_3$ (g) is _____ kJ/mol.

(A) -1108.7 B) +1108.7 C) -606.2 D) +606.2 E) -1,005

Thermodynamic Quantities for Selected Substances at 298.15 K (25 °C)

Substance	$\Delta H°_f$ (kJ/mol)	$\Delta G°_f$ (kJ/mol)	S (J/K-mol)
Chlorine			
Cl_2 (g)	0	0	222.96
Cl^- (aq)	-167.2	-131.2	56.5
Oxygen			
O_2 (g)	0	0	205.0
H_2O (l)	-285.83	-237.13	69.91
Phosphorus			
P_2 (g)	144.3	103.7	218.1
PCl_3 (g)	-288.1	-269.6	311.7
$POCl_3$ (g)	-542.2	-502.5	325

$\Delta G°_{rxn} = \sum n G°_f$ pdts $- \sum m G°_f$ reac

$= [2\, G°_{f\,POCl_3}] - [1\, G°_{f\,P_2}]$

$= \left[2 \times -502.5\frac{kJ}{mol}\right] - \left[1 \times \frac{103.7\,kJ}{mol}\right]$

$\Delta G° = -1108.7$ kJ/mol

13) Consider the reaction: **NH_3 (g) + HCl (g) → NH_4Cl (s)**
Given the following table of thermodynamic data at 298 K, the value of K for the reaction at 25 °C is _____.

Substance	$\Delta H_f°$ (kJ/mol)	S° (J/K · mol)
NH_3 (g)	-46.19	192.5
HCl (g)	-92.30	186.69
NH_4Cl (s)	-314.4	94.6

A) 150 (B) 9.3×10^{15} C) 8.4×10^{4} D) 1.1×10^{-16} E) 1.4×10^{8}

$\Delta G° = -RT\ln K$

$\Delta G° = \Delta H° - T\Delta S°$

$\Delta G° = \left(-175.9\frac{kJ}{mol}\right) - (298\,K)\left(-0.2846\frac{kJ}{mol \cdot K}\right)$

$\Delta G° = -91.10$ kJ/mol

$\Delta H° = \sum n H°_f$ pdts $- \sum m H°_f$ reactants

$= [1 H°_{f\,NH_4Cl}] - [(1 H°_{f\,NH_3}) + (1 H°_{f\,HCl})]$

$= \left[1 \times -314.4\frac{kJ}{mol}\right] - \left[\left(1 \times -46.19\frac{kJ}{mol}\right) + \left(1 \times -92.30\frac{kJ}{mol}\right)\right]$

$\Delta H° = -175.9$ kJ/mol

$\Delta S° = \sum n S°$ pdts $- \sum m S°$ reactants

$= [1 S°_{NH_4Cl}] - [(1 S°_{NH_3}) + (1 S°_{HCl})]$

$= \left[1 \times 94.6\frac{J}{mol \cdot K}\right] - \left[\left(1 \times 192.5\frac{J}{mol \cdot K}\right) + \left(1 \times 186.69\frac{J}{mol \cdot K}\right)\right]$

$\Delta S° = -284.6\,J/mol\,K \times \frac{1\,kJ}{1000\,J} = -0.2846\,kJ/mol \cdot K$

$K = e^{-(\Delta G°/RT)}$

$K = e^{-\left(\frac{-91.10\,kJ}{mol} \times \frac{1000\,J}{1\,kJ} \times \frac{mol \cdot K}{8.314\,J} \times \frac{1}{298\,K}\right)}$

$K = e^{-(-36.77)} = e^{+36.77}$

$K = 9.3 \times 10^{15}$